Answers to Chapter 3 Review Questions

(Student textbook pages 199-203)

1. b
2. a
3. c
4. a
5. e
6. d
7. a
8. a
9. b
10. d
11. e
12. a
13. c
14. d
15. In both models atoms were spherical and solid (no empty space) and were neutral particles. As well, in both models, atoms were considered the building blocks of all matter.
Thomson's model contained internal structures and discrete structures that were smaller than the atom itself. Specifically, Thomson's model had negatively charged subatomic particles, now known as electrons, embedded in a continuous mass of positively charged matter.
Dalton's atom was the smallest possible particle of matter. It was indivisible, meaning that it could not be broken down into smaller particles.
16. Thomson's and Rutherford's models of the atom were spherical and contained a sub-atomic particle believed to be the same particle in all matter (the electron).
Both models included equal amounts of positively and negatively charged matter to create an overall neutral atom.
In Rutherford's model, the positively charged matter was contained in a very small region at the centre of the atom which he called the nucleus. Most of the matter of the atom was contained within this nucleus. Negatively charged electrons orbited the nucleus. Most of the atoms consisted of empty space.
Thomson's model suggested that the positive and negatively charged matter was evenly distributed and no empty space existed within the atom.
17. In all atomic models, the atom is the building block of

all matter, atoms are neutrally charged, and atoms are spherical.

18. The speed of all electromagnetic radiation, in a vacuum, is the same, 3×10^8 m/s.

19. Bohr's electron moves in a circular pathway a fixed distance from the nucleus—its position is known. The electron in the quantum mechanical model moves constantly within a region of space that is described by using wave equations. Its exact position cannot be known for certain.

20. By using the basic laws of physics and Planck's concepts, Bohr was able to calculate the energy and radii of the allowed orbits for the hydrogen atom.

21. Groups 1 and 2 metals have low electron affinity and low ionization energies and therefore lose electrons easily, forming positive ions. These positive ions are then attracted to negative ions to form ionic compounds.

22. There are three equal-energy *p* orbitals. According to Hund's rule, single electrons with the same spin must occupy each equal-energy orbital before additional electrons with opposite spins can occupy the same orbitals. When the three electrons in the *p* orbitals of nitrogen have occupied the *p* orbitals, there are no additional electrons to pair with the single electrons in the *p* orbitals.

23. The exact position of an electron in an atom cannot be known with certainty. The energy of electrons in any electron shell and the relative size of the orbital are known with certainty as defined by the principal quantum number *n*.

24. Quantum numbers provide information about the energy, size, shape, spatial orientation, and the relative direction of the axis of the electron of the orbital.

25. The principle quantum number provides information about the energy and size of the orbital.

26. The electron in a higher energy level would have more energy to overcome the attractive pull of the positive charge of the nucleus and therefore be able to venture farther away from it.

27. The Pauli exclusion principle states that only two electrons of opposite spin could occupy an orbital. Each electron in an electron pair within the same orbital are identical but have opposite spins—the spin

quantum number is the same number for any electron, but the sign is either + (spin up) or - (spin down).

28. Hund's rule states that the lowest energy state for an atom has the maximum number of unpaired electrons allowed by the Pauli exclusion principle in a given energy sublevel. This rule can also be stated as single electrons with the same spin must occupy each equal-energy orbital before additional electrons with opposite spins can occupy the same orbitals. This rule determines the order of the filling of orbitals.

29. An atom's chemical properties are mainly associated with its ground state electron configuration. As well, there are a tremendous number of possible excited states.

30. The atomic radius decreases across a period (from left to right) and increases down a group. The reason that it decreases across a period is because the amount of charge in the nucleus becomes greater across a period. Thus the attractive force on the electrons becomes greater, pulling the electrons closer to the nucleus. The radius increases down a group because each successive group has one more energy level, or shell, than the previous one and thus electrons are farther from the nucleus.

31. None of the experiments described in the book would have remained the same because the tests depend on the charge carried by the particles.

32. The element described, would be in Period 6, *d* block, and Group 12.

33. $[Kr]5s^24d^{10}5p^3$

34. 3*s*, 3*d*, 4*p*, 5*s*, 5*d*

35. The highest possible value of l is 3 (because there are electrons in a *d* orbital). Therefore, m_l can be anything from .3 to +3.

36. nickel, Ni

37. aluminum, Al

38. tellurium, Te

39. chlorine, Cl

40. ruthenium, Ru

41. oxygen, O

42. barium, Ba

43. In these cases, the *d* electrons do not typically participate in chemical reactions. As well, all *p*-block elements in Period 4 and above have 10 *d* electrons.

44. Boron will be larger than fluorine.

45. Magnesium will be larger than silicon.
46. Calcium will be larger than selenium.
47. a. Electron configuration: $[Rn]7s^{2}5f^{14}6d^{10}7p^{6}$
b. Quantum number for last electron: $n = 7$, $l = 1$,
$m_l = 1$, $m_s = -\frac{1}{2}$
c. Physical state at room temperature: gas
d. Reactivity: inert, just like all other noble gases
48. a. Na, Si, Ar: First ionization energy tends to increase across a period. Sodium, silicon, and argon are all in Period 3.
b. Ba, Ca, Mg: First ionization energy tends to decrease down a group. Barium, calcium, and magnesium are all in Group 2.
c. Li, Be, He: Helium would have the highest first ionization energy because it is the farthest right in Period 1. Lithium would have the lowest first ionization energy because it is to the left of beryllium in Period 2.
49. a. Oxygen: There are eight electrons in the neutral atom, thus there must be eight protons in the nucleus. Oxygen has the atomic number eight.
b. The electron configuration indicates the values of the quantum numbers n and l.
c. The electron configuration does not provide information about the magnetic and spin quantum numbers, m_s and m_l.